Table of Contents

Lesson Starter

- CCl_4 $MgCl_2$
- Guess the name of each of the above compounds based on the formulas written.
- What kind of information can you discern from the formulas?
- Guess which of the compounds represented is molecular and which is ionic.
- Chemical formulas form the basis of the language of chemistry and reveal much information about the substances they represent.

Objectives

- **Explain** the significance of a chemical formula.
- **Determine** the formula of an ionic compound formed between two given ions.
- **Name** an ionic compound given its formula.
- Using prefixes, **name** a binary molecular compound from its formula.
- **Write** the formula of a binary molecular compound given its name.

Significance of a Chemical Formula

- A chemical formula indicates the relative number of atoms of each kind in a chemical compound.
- For a molecular compound, the chemical formula reveals the number of atoms of each element contained in a single molecule of the compound.
  - example: octane — C_8H_{18}

The subscript after the C indicates that there are 8 carbon atoms in the molecule.

The subscript after the H indicates that there are 18 hydrogen atoms in the molecule.

Significance of a Chemical Formula, *continued*

- The chemical formula for an ionic compound represents one formula unit—the simplest ratio of the compound's positive ions (cations) and its negative ions (anions).
  - example: aluminum sulfate — $Al_2(SO_4)_3$
  - Parentheses surround the polyatomic ion SO_4^{2-} to identify it as a unit. The subscript *3* refers to the unit.
  - Note also that there is no subscript for sulfur: when there is no subscript next to an atom, the subscript is understood to be 1.

Monatomic Ions

- Many main-group elements can lose or gain electrons to form ions.
- Ions formed form a single atom are known as monatomic ions.
  - example: To gain a noble-gas electron configuration, nitrogen gains three electrons to form N^{3-} ions.
- Some main-group elements tend to form covalent bonds instead of forming ions.
  - examples: carbon and silicon

Monatomic Ions, *continued*

Naming Monatomic Ions

- Monatomic cations are identified simply by the element's name.
  - examples:
    - K^+ is called the potassium cation
    - Mg^{2+} is called the magnesium cation
- For monatomic anions, the ending of the element's name is dropped, and the ending -ide is added to the root name.
  - examples:
    - F^- is called the fluoride anion
    - N^{3-} is called the nitride anion

Common Monatomic Ions

Main-group elements

1+		2+		3+	
lithium	Li^+	beryllium	Be^{2+}		
sodium	Na^+	magnesium	Mg^{2+}	aluminum	Al^{3+}
potassium	K^+	calcium	Ca^{2+}		
rubidium	Rb^+	strontium	Sr^{2+}		
cesium	Cs^+	barium	Ba^{2+}		
1–		**2–**		**3–**	
fluoride	F^-	oxide	O^{2-}	nitride	N^{3-}
chloride	Cl^-	sulfide	S^{2-}	phosphide	P^{3-}
bromide	Br^-				
iodide	I^-				

Common Monatomic Ions

***d*-Block elements**

1+		2+	
copper(I)	Cu^+	vanadium(II)	V^{2+}
silver	Ag^+	chromium(II)	Cr^{2+}
		manganese(II)	Mn^{2+}
		iron(II)	Fe^{2+}
		cobalt(II)	Co^{2+}
		nickel(II)	Ni^{2+}
		copper(II)	Cu^{2+}
		zinc	Zn^{2+}
		cadmium	Cd^{2+}
		tin(II)	Sn^{2+}
		mercury(II)	Hg^{2+}
		lead(II)	Pb^{2+}

3+		4+	
vanadium(III)	V^{3+}	vanadium(IV)	V^{4+}
chromium(III)	Cr^{3+}	tin(IV)	Sn^{4+}
iron(III)	Fe^{3+}	lead(IV)	Pb^{4+}
cobalt(III)	Co^{3+}		

Binary Ionic Compounds

- Compounds composed of two elements are known as **binary compounds.**
- In a binary ionic compound, the total numbers of positive charges and negative charges must be equal.
- The formula for a binary ionic compound can be written given the identities of the compound's ions.
  - example: magnesium bromide
    Ions combined: Mg^{2+}, Br^-, Br^-
    Chemical formula: $MgBr_2$

Binary Ionic Compounds, *continued*

- A general rule to use when determining the formula for a binary ionic compound is "crossing over" to balance charges between ions.
  - example: aluminum oxide

    1) Write the symbols for the ions.

    $$Al^{3+}\ \ O^{2-}$$

    2) Cross over the charges by using the absolute value of each ion's charge as the subscript for the other ion.

    $Al_2^{3+}\ O_3^{2-}$

Binary Ionic Compounds, *continued*

- example: aluminum oxide, *continued*

  $Al_2^{3+}\ O_3^{2-}$

3) Check the combined positive and negative charges to see if they are equal.

$$(2 \times 3+) + (3 \times 2–) = 0$$

The correct formula is Al_2O_3

Writing the Formula of an Ionic Compound

Follow the following steps when writing the formula of a binary ionic compound, such as iron(III) oxide.

- Write the symbol and charges for the cation and anion.
  The roman numeral indicates which cation iron forms.
  symbol for iron(III): Fe^{3+} symbol for oxide: O^{2-}
- Write the symbols for the ions side by side, beginning with the cation.
  $Fe^{3+}O^{2-}$
- To determine how to get a neutral compound, look for the lowest common multiple of the charges on the ions. The lowest common multiple of 3 and 2 is 6. Therefore, the formula should indicate six positive charges and six negative charges.

For six positive charges, you need two Fe^{3+} ions because $2 \times 3+ = 6+$.

For six negative charges, you need three O^{2-} ions because $3 \times 2- = 6-$.

Therefore the ratio of Fe^{3+} to O^{2-} is 2Fe:3O. The formula is written as follows.

Fe_2O_3

Naming Binary Ionic Compounds

- The **nomenclature,** or naming system, or binary ionic compounds involves combining the names of the compound's positive and negative ions.
- The name of the cation is given first, followed by the name of the anion:
  - example: Al_2O_3 — aluminum oxide
- For most simple ionic compounds, the ratio of the ions is not given in the compound's name, because it is understood based on the relative charges of the compound's ions.

Naming Binary Ionic Compounds, *continued*

Sample Problem A

Write the formulas for the binary ionic compounds formed between the following elements:

a. zinc and iodine

b. zinc and sulfur

Naming Binary Ionic Compounds, *continued*

The Stock System of Nomenclature

- Some elements such as iron, form two or more cations with different charges.
- To distinguish the ions formed by such elements, scientists use the Stock system of nomenclature.
- The system uses a Roman numeral to indicate an ion's charge.
  - examples: Fe^{2+} iron(II)
    Fe^{3+} iron(III)

Naming Binary Ionic Compounds, *continued*

The Stock System of Nomenclature, *continued*

Sample Problem B

Write the formula and give the name for the compound formed by the ions Cr^{3+} and F^{-}.

Naming Binary Ionic Compounds, *continued*

Compounds Containing Polyatomic Ions

- Many common polyatomic ions are **oxyanions**—polyatomic ions that contain oxygen.
- Some elements can combine with oxygen to form more than one type of oxyanion.
  - example: nitrogen can form NO_3^- or NO_2^-.
    - The name of the ion with the greater number of oxygen atoms ends in *-ate.* The name of the ion with the smaller number of oxygen atoms ends in *-ite.*

NO_3^-	NO_2^-
nitrate	**nitrite**

Naming Binary Ionic Compounds, *continued*
Compounds Containing Polyatomic Ions, *continued*

- Some elements can form more than two types of oxyanions.
  - example: chlorine can form ClO^-, ClO_2^-, ClO_3^-, or ClO_4^-
    - In this case, an anion that has one fewer oxygen atom than the *-ite* anion has is given the prefix *hypo-*.
    - An anion that has one more oxygen atom than the *-ate* anion has is given the prefix *per-*.

ClO^-	ClO_2^-	ClO_3^-	ClO_4^-
hypochlorite	**chlorite**	**chlorate**	**perchlorate**

Naming Compounds with Polyatomic Ions

Follow these steps when naming an ionic compound that contains one or more polyatomic ions, such as K_2CO_3.

- Name the cation. Recall that a cation is simply the name of the element. In this formula, K is potassium that forms a singly charged cation, K^+, of the same name.
- Name the anion. Recall that salts are electrically neutral. Because there are two K^+ cations present in this salt, these two positive charges must be balanced by two negative charges. Therefore, the polyatomic anion in this salt must be CO_3^{2-}. You may find it helpful to think of the formula as follows, although it is not written this way.

$$(K^+)_2(CO_3^{2-})$$

The CO_3^{2-} polyatomic ion is called carbonate.

- Name the salt. Recall that the name of a salt is just the names of the cation and anion. The salt K_2CO_3 is potassium carbonate.

Naming Binary Ionic Compounds, *continued*
Compounds Containing Polyatomic Ions, *continued*

Sample Problem C

Write the formula for tin(IV) sulfate.

Naming Binary Molecular Compounds

- Unlike ionic compounds, molecular compounds are composed of individual covalently bonded units, or molecules.
- As with ionic compounds, there is also a Stock system for naming molecular compounds.
- The old system of naming molecular compounds is based on the use of prefixes.
  - examples: CCl_4 — carbon *tetra*chloride (*tetra-* = 4)
    CO — carbon *mono*xide (*mon-* = 1)
    CO_2 — carbon *di*oxide (*di-* = 2)

Prefixes for Naming Covalent Compounds

Prefix	Number of Atoms	Example	Name
mono-	1	CO	carbon monoxide
di-	2	SiO_2	silicon dioxide
tri-	3	SO_3	sulfur trioxide
tetra-	4	SCl_4	sulfur tetrachloride
penta-	5	$SbCl_5$	antimony pentachloride
Prefix	**Number of Atoms**	**Example**	**Name**
hexa-	6	CeB_6	cerium hexaboride
hepta-	7	IF_7	iodine heptafluoride
octa-	8	Np_3O_8	trineptunium octoxide
nona-	9	I_4O_9	tetraiodine nonoxide
deca-	10	S_2F_{10}	disulfur decafluoride

Naming Binary Molecular Compounds, *continued*

Sample Problem D

a. Give the name for As_2O_5.

b. Write the formula for oxygen difluoride.

Covalent-Network Compounds

- Some covalent compounds do not consist of individual molecules.
- Instead, each atom is joined to all its neighbors in a covalently bonded, three-dimensional network.
- Subscripts in a formula for covalent-network compound indicate smallest whole-number ratios of the atoms in the compound.
  - examples: SiC, silicon carbide
    SiO_2, silicon dioxide
    Si_3N_4, trisilicon tetranitride.

Acids and Salts

- An *acid* is a certain type of molecular compound. Most acids used in the laboratory are either binary acids or oxyacids.
  - *Binary acids* are acids that consist of two elements, usually hydrogen and a halogen.
  - *Oxyacids* are acids that contain hydrogen, oxygen, and a third element (usually a nonmetal).

Acids and Salts, *continued*

- In the laboratory, the term *acid* usually refers to a solution in water of an acid compound rather than the acid itself.
  - example: *hydrochloric acid* refers to a water solution of the molecular compound hydrogen chloride, HCl
- Many polyatomic ions are produced by the loss of hydrogen ions from oxyacids.
  - examples:

sulfuric acid	H_2SO_4	sulfate	SO_4^{2-}
nitric acid	HNO_3	nitrate	NO_3^-
phosphoric acid	H_3PO_4	phosphate	PO_4^{3-}

Acids and Salts, *continued*

- An ionic compound composed of a cation and the anion from an acid is often referred to as a **salt.**
  - examples:
    - Table salt, NaCl, contains the anion from hydrochloric acid, HCl.
    - Calcium sulfate, $CaSO_4$, is a salt containing the anion from sulfuric acid, H_2SO_4.
    - The bicarbonate ion, HCO_3^-, comes from carbonic acid, H_2CO_3.

Lesson Starter

- It is possible to determine the charge of an ion in an ionic compound given the charges of the other ions present in the compound.
- Determine the charge on the bromide ion in the compound NaBr given that Na^+ has a 1+ charge.
- Answer: The total charge is 0, so Br^- must have a charge of 1– in order to balance the 1+ charge of Na^+.

Lesson Starter, *continued*

- Numbers called oxidation numbers can be assigned to atoms in order to keep track of electron distributions in molecular as well as ionic compounds.

Objectives

- **List** the rules for assigning oxidation numbers.
- **Give** the oxidation number for each element in the formula of a chemical compound.
- **Name** binary molecular compounds using oxidation numbers and the Stock system.

Oxidation Numbers

- The charges on the ions in an ionic compound reflect the electron distribution of the compound.
- In order to indicate the general distribution of electrons among the bonded atoms in a molecular compound or a polyatomic ion, **oxidation numbers** are assigned to the atoms composing the compound or ion.
- Unlike ionic charges, oxidation numbers do not have an exact physical meaning: rather, they serve as useful "bookkeeping" devices to help keep track of electrons.

Assigning Oxidation Numbers

- In general when assigning oxidation numbers, shared electrons are assumed to "belong" to the more electronegative atom in each bond.
- More-specific rules are provided by the following guidelines.
  1. The atoms in a pure element have an oxidation number of zero.

     examples: all atoms in sodium, Na, oxygen, O_2, phosphorus, P_4, and sulfur, S_8, have oxidation numbers of zero.

Assigning Oxidation Numbers, *continued*

2. The more-electronegative element in a binary compound is assigned a negative number equal to the charge it would have as an anion. Likewise for the less-electronegative element.
3. Fluorine has an oxidation number of –1 in all of its compounds because it is the most electronegative element.

Assigning Oxidation Numbers, *continued*

4. Oxygen usually has an oxidation number of –2.

   Exceptions:

   - In peroxides, such as H_2O_2, oxygen's oxidation number is –1.
   - In compounds with fluorine, such as OF_2, oxygen's oxidation number is +2.
5. Hydrogen has an oxidation number of +1 in all compounds containing elements that are more electronegative than it; it has an oxidation number of –1 with metals.

Assigning Oxidation Numbers, *continued*

6. The algebraic sum of the oxidation numbers of all atoms in an neutral compound is equal to zero.
7. The algebraic sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the charge of the ion.
8. Although rules 1 through 7 apply to covalently bonded atoms, oxidation numbers can also be applied to atoms in ionic compounds similarly.

Assigning Oxidation Numbers, *continued*

Sample Problem E

Assign oxidation numbers to each atom in the following compounds or ions:

a. UF_6

b. H_2SO_4

c. ClO_3^-

Using Oxidation Numbers for Formulas and Names

- As shown in the table in the next slide, many nonmetals can have more than one oxidation number.
- These numbers can sometimes be used in the same manner as ionic charges to determine formulas.
  - example: What is the formula of a binary compound formed between sulfur and oxygen?

    From the common +4 and +6 oxidation states of sulfur, you could predict that sulfur might form SO_2 or SO_3.

    Both are known compounds.

Common Oxidation States of Nonmetals

Group 14	carbon	−4, +2, +4
Group 15	nitrogen	−3, +1, +2, +3, +4, +5
	phosphorus	−3, +3, +5
Group 16	sulfur	−2, +4, +6
Group 17	chlorine	−1, +1, +3, +5, +7
	bromine	−1, +1, +3, +5, +7
	iodine	−1, +1, +3, +5, +7

In addition to the values shown, atoms of each element in its pure state are assigned an oxidation number of zero.

Using Oxidation Numbers for Formulas and Names, *continued*

- Using oxidation numbers, the Stock system, introduced in the previous section for naming ionic compounds, can be used as an alternative to the prefix system for naming binary molecular compounds.

	Prefix system	Stock system
PCl_3	phosphorus trichloride	phosphorus(III) chloride
PCl_5	phosphorus pentachloride	phosphorus(V) chloride
N_2O	dinitrogen monoxide	nitrogen(I) oxide
NO	nitrogen monoxide	nitrogen(II) oxide
Mo_2O_3	dimolybdenum trioxide	molybdenum(III) oxide

Lesson Starter

- The chemical formula for water is H_2O.
- How many atoms of hydrogen and oxygen are there in one water molecule?
- How might you calculate the mass of a water molecule, given the atomic masses of hydrogen and oxygen?
- In this section, you will learn how to carry out these and other calculations for any compound.

Objectives

- **Calculate** the formula mass or molar mass of any given compound.
- **Use** molar mass to convert between mass in grams and amount in moles of a chemical compound.
- **Calculate** the number of molecules, formula units, or ions in a given molar amount of a chemical compound.
- **Calculate** the percentage composition of a given chemical compound.

Chapter 7 Section 3 Using Chemical Formulas

- A chemical formula indicates:
  - the elements present in a compound
  - the relative number of atoms or ions of each element present in a compound
- Chemical formulas also allow chemists to calculate a number of other characteristic values for a compound:
  - *formula mass*
  - *molar mass*
  - *percentage composition*

Chapter 7 Section 3 Using Chemical Formulas

Formula Masses

- The **formula mass** of any molecule, formula unit, or ion is the sum of the average atomic masses of all atoms represented in its formula.
  - example: formula mass of water, H_2O
    average atomic mass of H: 1.01 amu
    average atomic mass of O: 16.00 amu

$$2\ \cancel{\text{H atoms}} \times \frac{1.01\ \text{amu}}{\cancel{\text{H atom}}} = 2.02\ \text{amu}$$

$$1\ \cancel{\text{O atom}} \times \frac{16.00\ \text{amu}}{\cancel{\text{O atom}}} = 16.00\ \text{amu}$$

average mass of H_2O molecule: 18.02 amu

Chapter 7 Section 3 Using Chemical Formulas

Formula Masses

- The mass of a water molecule can be referred to as a *molecular mass.*
- The mass of one formula unit of an ionic compound, such as NaCl, is not a molecular mass.
- The mass of any unit represented by a chemical formula (H_2O, NaCl) can be referred to as the formula mass.

Chapter 7 Section 3 Using Chemical Formulas

Formula Masses, *continued*

Sample Problem F

Find the formula mass of potassium chlorate, $KClO_3$.

Chapter 7 Section 3 Using Chemical Formulas

Molar Masses

- The molar mass of a substance is equal to the mass in grams of one mole, or approximately 6.022×10^{23} particles, of the substance.
  - example: the molar mass of pure calcium, Ca, is 40.08 g/mol because one mole of calcium atoms has a mass of 40.08 g.
- The molar mass of a compound is calculated by adding the masses of the elements present in a mole of the molecules or formula units that make up the compound.

Chapter 7 Section 3 Using Chemical Formulas

Molar Masses, *continued*

- One mole of water molecules contains exactly two moles of H atoms and one mole of O atoms. The molar mass of water is calculated as follows.

$$2\ \cancel{\text{mol H}} \times \frac{1.01\ \text{g H}}{\cancel{\text{mol H}}} = 2.02\ \text{g H}$$

$$1\ \cancel{\text{mol O}} \times \frac{16.00\ \text{g O}}{\cancel{\text{mol O}}} = 16.00\ \text{g O}$$

molar mass of H_2O molecule: 18.02 g/mol

- A compound's molar mass is numerically equal to its formula mass.

Molar Masses, *continued*

Sample Problem G

What is the molar mass of barium nitrate, $Ba(NO_3)_2$?

Molar Mass as a Conversion Factor

- The molar mass of a compound can be used as a conversion factor to relate an amount in moles to a mass in grams for a given substance.
- To convert moles to grams, multiply the amount in moles by the molar mass:

$$\text{Amount in moles} \times \text{molar mass (g/mol)} = \text{mass in grams}$$

Molar Mass as a Conversion Factor, *continued*

Sample Problem H

What is the mass in grams of 2.50 mol of oxygen gas?

Molar Mass as a Conversion Factor, *continued*

Sample Problem I

Ibuprofen, $C_{13}H_{18}O_2$, is the active ingredient in many nonprescription pain relievers. Its molar mass is 206.31 g/mol.

a. If the tablets in a bottle contain a total of 33 g of ibuprofen, how many moles of ibuprofen are in the bottle?

b. How many molecules of ibuprofen are in the bottle?

c. What is the total mass in grams of carbon in 33 g of ibuprofen?

Percentage Composition

- It is often useful to know the percentage by mass of a particular element in a chemical compound.
- To find the mass percentage of an element in a compound, the following equation can be used.

$$\frac{\text{mass of element in sample of compound}}{\text{mass of sample of compound}} \times 100 = \text{\% element in compound}$$

- The mass percentage of an element in a compound is the same regardless of the sample's size.

Percentage Composition, *continued*

- The percentage of an element in a compound can be calculated by determining how many grams of the element are present in one mole of the compound.

$$\frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100 = \text{\% element in compound}$$

- The percentage by mass of each element in a compound is known as the **percentage composition** of the compound.

Chapter 7 **Section 3 Using Chemical Formulas**

Percentage Composition, *continued*

Sample Problem J

Find the percentage composition of copper(I) sulfide, Cu_2S.

Chapter 7 **Section 4 Determining Chemical Formulas**

Lesson Starter

- Compare and contrast models of the molecules NO_2 and N_2O_4.
- The numbers of atoms in the molecules differ, but the ratio of N atoms to O atoms for each molecule is the same.

Chapter 7 **Section 4 Determining Chemical Formulas**

Objectives

- **Define** *empirical formula,* and explain how the term applies to ionic and molecular compounds.
- **Determine** an empirical formula from either a percentage or a mass composition.
- **Explain** the relationship between the empirical formula and the molecular formula of a given compound.
- **Determine** a molecular formula from an empirical formula.

Chapter 7 **Section 4 Determining Chemical Formulas**

- An **empirical formula** consists of the symbols for the elements combined in a compound, with subscripts showing the smallest whole-number mole ratio of the different atoms in the compound.
- For an ionic compound, the formula unit is usually the compound's empirical formula.
- For a molecular compound, however, the empirical formula does not necessarily indicate the actual numbers of atoms present in each molecule.
  - example: the empirical formula of the gas diborane is BH_3, but the molecular formula is B_2H_6.

Chapter 7 **Section 4 Determining Chemical Formulas**

Calculation of Empirical Formulas

- To determine a compound's empirical formula from its percentage composition, begin by converting percentage composition to a mass composition.
  - Assume that you have a 100.0 g sample of the compound.
  - Then calculate the amount of each element in the sample.
  - example: diborane
    - The percentage composition is 78.1% B and 21.9% H.
    - Therefore, 100.0 g of diborane contains 78.1 g of B and 21.9 g of H.

Chapter 7 **Section 4 Determining Chemical Formulas**

Calculation of Empirical Formulas, *continued*

- Next, the mass composition of each element is converted to a composition in moles by dividing by the appropriate molar mass.

$$78.1\ \text{g B} \times \frac{1\ \text{mol B}}{10.81\ \text{g B}} = 7.22\ \text{mol B}$$

$$21.9\ \text{g H} \times \frac{1\ \text{mol H}}{1.01\ \text{g H}} = 21.7\ \text{mol H}$$

- These values give a mole ratio of 7.22 mol B to 21.7 mol H.

Calculation of Empirical Formulas, *continued*

- To find the smallest whole number ratio, divide each number of moles by the smallest number in the existing ratio.

$$\frac{7.22 \text{ mol B}}{7.22} : \frac{21.7 \text{ mol H}}{7.22} = 1 \text{ mol B} : 3.01 \text{ mol H}$$

- Because of rounding or experimental error, a compound's mole ratio sometimes consists of numbers close to whole numbers instead of exact whole numbers.
  - In this case, the differences from whole numbers may be ignored and the nearest whole number taken.

Calculation of Empirical Formulas, *continued*

Sample Problem L

Quantitative analysis shows that a compound contains 32.38% sodium, 22.65% sulfur, and 44.99% oxygen. Find the empirical formula of this compound.

Calculation of Molecular Formulas

- The *empirical formula* contains the smallest possible whole numbers that describe the atomic ratio.
- The *molecular formula* is the actual formula of a molecular compound.
- An empirical formula may or may not be a correct molecular formula.
- The relationship between a compound's empirical formula and its molecular formula can be written as follows.

x(empirical formula) = molecular formula

Calculation of Molecular Formulas, *continued*

- The formula masses have a similar relationship.

  x(empirical formula mass) = molecular formula mass

- To determine the molecular formula of a compound, you must know the compound's formula mass.
  - Dividing the experimental formula mass by the empirical formula mass gives the value of x.
- A compound's molecular formula mass is numerically equal to its molar mass, so a compound's molecular formula can also be found given the compound's empirical formula and its molar mass.

Calculation of Molecular Formulas, *continued*

Sample Problem N

In Sample Problem M, the empirical formula of a compound of phosphorus and oxygen was found to be P_2O_5. Experimentation shows that the molar mass of this compound is 283.89 g/mol. What is the compound's molecular formula?